John E. McMurry • Robert C. Fay

# General Chemistry: Atoms First

## Chapter 7
## Reactions in Aqueous Solution

**Lecture Notes**

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# Reactions in Solution

1. Precipitation Reactions
2. Neutralization Reactions
3. Redox Reactions

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# Some Ways That Chemical Reactions Occur

**Precipitation Reaction**: A process in which soluble ionic reactants yield an insoluble solid product that falls out of solution.

$$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow 2KNO_3(aq) + PbI_2(s)$$

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# Some Ways That Chemical Reactions Occur

**Acid-Base Neutralization Reactions**: Processes in which an acid reacts with a base to yield water plus an ionic compound called a *salt*.

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$

# Some Ways That Chemical Reactions Occur

**Oxidation-Reduction (Redox) Reactions**: Processes in which one or more electrons are transferred between reaction partners (atoms, molecules, or ions).

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

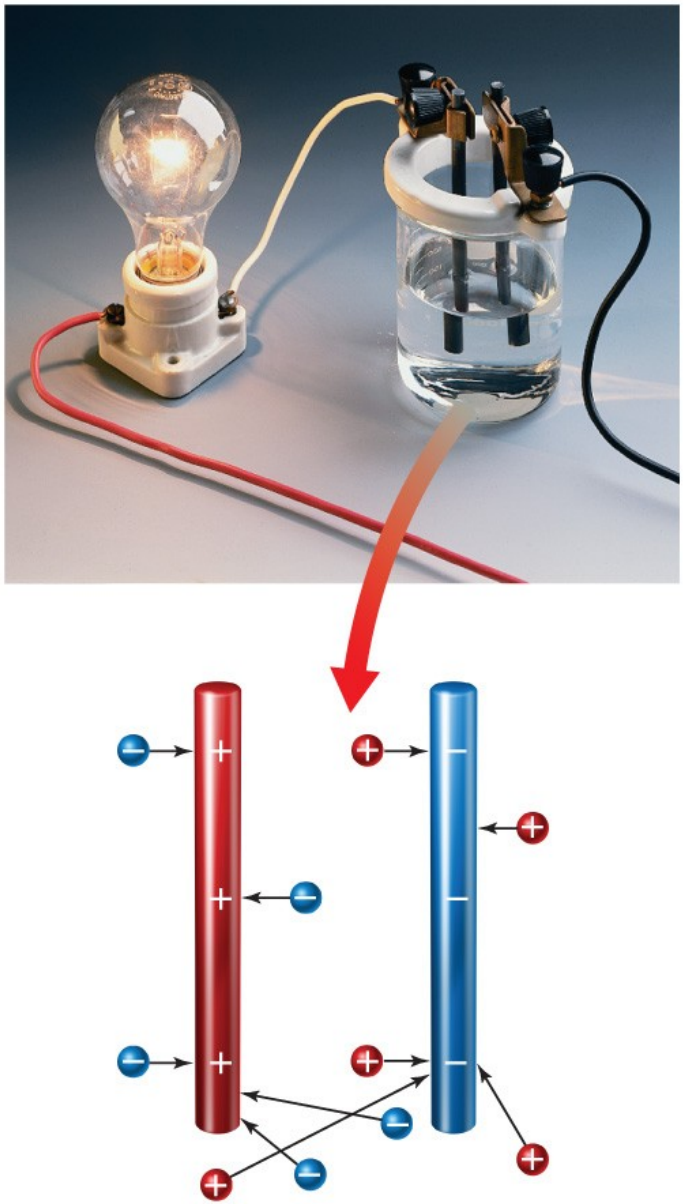
# Electrolytes in Aqueous Solution

**Electrolytes**: Substances which dissolve in water to produce conducting solutions of ions.

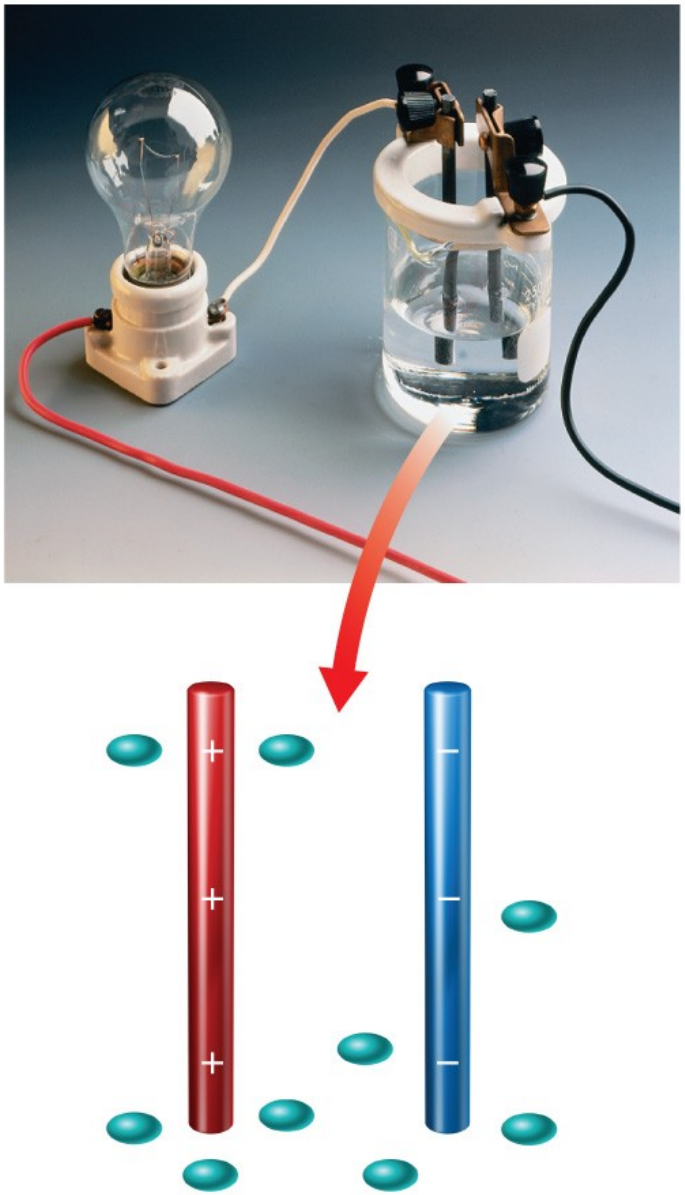
$$NaCl(s) \xrightarrow{H_2O} Na^{1+}(aq) + Cl^{1-}(aq)$$

**Nonelectrolytes**: Substances which do not produce ions in aqueous solutions.

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}(aq)$$



A solution of NaCl conducts electricity because of the movement of charged particles (ions), thereby completing the circuit and allowing the bulb to light.



A solution of sucrose does not conduct electricity or complete the circuit because it contains no charged particles. The bulb therefore remains dark.

# Electrolytes in Aqueous Solution

**Strong Electrolytes**: Compounds that dissociate to a *large* extent into ions when dissolved in water.

$$HCl(g) \xrightarrow{H_2O} H^{1+}(aq) + Cl^{1-}(aq)$$

**Weak Electrolytes**: Compounds that dissociate to a *small* extent into ions when dissolved in water.

$$CH_3CO_2H(aq) \rightleftharpoons H^{1+}(aq) + CH_3CO_2^{1-}(aq)$$

# Electrolytes in Aqueous Solution

**TABLE 7.1** Electrolyte Classification of Some Common Substances

| Strong Electrolytes | Weak Electrolytes | Nonelectrolytes |
|---|---|---|
| HCl, HBr, HI | $CH_3CO_2H$ | $H_2O$ |
| $HClO_4$ | HF | $CH_3OH$ (methyl alcohol) |
| $HNO_3$ | | $C_2H_5OH$ (ethyl alcohol) |
| $H_2SO_4$ | | $C_{12}H_{22}O_{11}$ (sucrose) |
| KBr | | Most compounds of carbon (organic compounds) |
| NaCl | | |
| NaOH, KOH | | |
| Other soluble ionic compounds | | |

***Strong* Acids**: hydrochloric acid, hydrobromic acid, hydroiodic acid, perchloric acid, nitric acid,

# Electrolytes in Aqueous Solution

**TABLE 7.1** Electrolyte Classification of Some Common Substances

| Strong Electrolytes | Weak Electrolytes | Nonelectrolytes |
|---|---|---|
| HCl, HBr, HI | $CH_3CO_2H$ | $H_2O$ |
| $HClO_4$ | HF | $CH_3OH$ (methyl alcohol) |
| $HNO_3$ | | $C_2H_5OH$ (ethyl alcohol) |
| $H_2SO_4$ | | $C_{12}H_{22}O_{11}$ (sucrose) |
| KBr | | Most compounds of carbon (organic compounds) |
| NaCl | | |
| NaOH, KOH | | |
| Other soluble ionic compounds | | |

**Ionic compounds**

# Electrolytes in Aqueous Solution

**TABLE 7.1** Electrolyte Classification of Some Common Substances

| Strong Electrolytes | Weak Electrolytes | Nonelectrolytes |
|---|---|---|
| HCl, HBr, HI | $CH_3CO_2H$ | $H_2O$ |
| $HClO_4$ | HF | $CH_3OH$ (methyl alcohol) |
| $HNO_3$ | | $C_2H_5OH$ (ethyl alcohol) |
| $H_2SO_4$ | | $C_{12}H_{22}O_{11}$ (sucrose) |
| KBr | | Most compounds of carbon (organic compounds) |
| NaCl | | |
| NaOH, KOH | | |
| Other soluble ionic compounds | | |

***Weak* acids**

# Electrolytes in Aqueous Solution

**TABLE 7.1** Electrolyte Classification of Some Common Substances

| Strong Electrolytes | Weak Electrolytes | Nonelectrolytes |
|---|---|---|
| HCl, HBr, HI | $CH_3CO_2H$ | $H_2O$ |
| $HClO_4$ | HF | $CH_3OH$ (methyl alcohol) |
| $HNO_3$ | | $C_2H_5OH$ (ethyl alcohol) |
| $H_2SO_4$ | | $C_{12}H_{22}O_{11}$ (sucrose) |
| KBr | | Most compounds of carbon (organic compounds) |
| NaCl | | |
| NaOH, KOH | | |
| Other soluble ionic compounds | | |

**Molecular compounds** (other than any strong or weak electrolytes)

# Electrolytes in Aqueous Solution

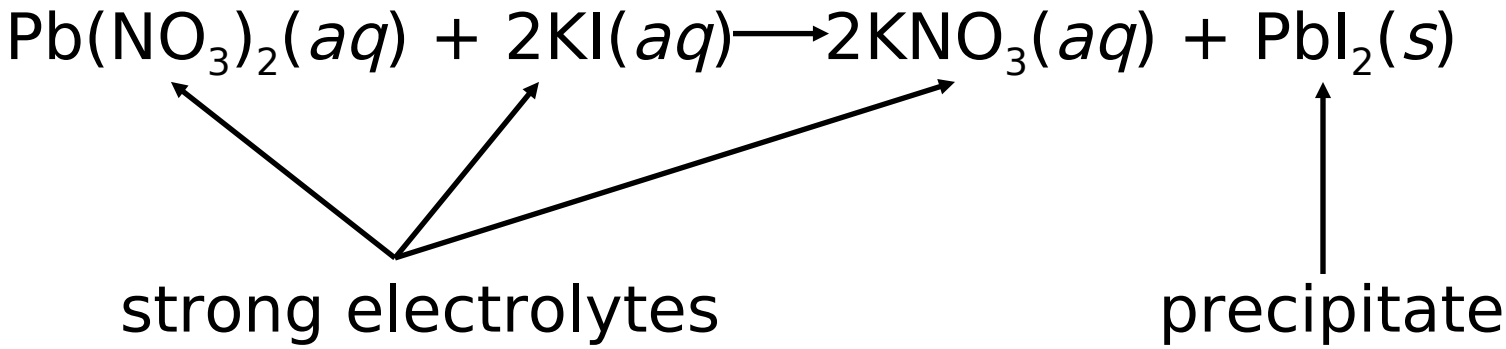
## Dissociation Equations

$$Na_2SO_4(s) \xrightarrow{H_2O} 2Na^{1+}(aq) + SO_4^{2-}(aq)$$

$$FeBr_3(s) \xrightarrow{H_2O} Fe^{3+}(aq) + 3Br^{1-}(aq)$$

# Aqueous Reactions and Net Ionic Equations

**Molecular Equation**: All substances in the chemical equation are written using their complete formulas as if they were *molecules*.



# Aqueous Reactions and Net Ionic Equations

**Ionic Equation**: All of the strong electrolytes are written as ions.

$$\overbrace{Pb(NO_3)_2(aq)}^{} \qquad \overbrace{2KI(aq)}^{}$$

$$\overbrace{Pb^{2+}(aq) + 2NO_3^{1-}(aq)}^{} + \overbrace{2K^{1+}(aq) + 2I^{1-}(aq)}^{}$$

$$\longrightarrow \overbrace{2K^{1+}(aq) + 2NO_3^{1-}(aq)}^{} + PbI_2(s)$$

$$\overbrace{2KNO_3(aq)}^{}$$

# Aqueous Reactions and Net Ionic Equations

**Spectator Ions**: Ions that undergo no change during the reaction and appear on both sides of the reaction arrow.

$$Pb^{2+}(aq) + \boxed{2NO_3^{1-}(aq)} + \boxed{2K^{1+}(aq)} + 2I^{1-}(aq) \longrightarrow \boxed{2K^{1+}(aq)} + \boxed{2NO_3^{1-}(aq)} + PbI_2(s)$$

# Aqueous Reactions and Net Ionic Equations

**Net Ionic Equation**: Only the ions undergoing change are shown.

$$Pb^{2+}(aq) + 2I^{1-}(aq) \longrightarrow PbI_2(s)$$

# Precipitation Reactions and Solubility Guidelines

**Solubility**: States how much of a compound will dissolve in a given amount of solvent at a given temperature.

# Precipitation Reactions and Solubility Guidelines

1. **A compound is probably soluble if it contains one of the following *cations*:**
   - Group 1A cation: $Li^{1+}$, $Na^{1+}$, $K^{1+}$, $Cs^{1+}$
   - Ammonium ion: $NH_4^{1+}$

2. **A compound is probably soluble if it contains one of the following *anions*:**
   - Halide: $Cl^{1-}$, $Br^{1-}$, $I^{1-}$
     - *except* $Ag^{1+}$, $Hg_2^{2+}$, and $Pb^{2+}$ compounds
   - Nitrate ($NO_3^{1-}$), perchlorate ($ClO_4^{1-}$), acetate ($CH_3CO_2^{1-}$), sulfate ($SO_4^{2-}$)
     - *except* $Ba^{2+}$, $Hg_2^{2+}$, and $Pb^{2+}$ sulfates

# Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and $Na_2CO_3$ are mixed.

# Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and $Na_2CO_3$ are mixed.

1. Write the unbalanced chemical formulas of the products (use proper ionic rules).

$$AgNO_3(aq) + Na_2CO_3(aq) \longrightarrow Ag_2CO_3 + NaNO_3$$

$$AB + CD \longrightarrow CB + AD$$

double replacement reactic

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# Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and $Na_2CO_3$ are mixed.

2. **Molecular Equation**: Balance the equation and predict the solubility of each possible product.

$$2AgNO_3(aq) + Na_2CO_3(aq) \longrightarrow Ag_2CO_3(s) + 2NaNO_3(aq)$$

Neither the cation nor the anion is in the solubility list.

Contains a group 1A cation.

# Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and $Na_2CO_3$ are mixed.

3. **Ionic Equation**: Dissociate the soluble ionic compounds.

$$\overbrace{2AgNO_3(aq)}^{2Ag^{1+}(aq) + 2NO_3^{1-}(aq)} \quad \overbrace{Na_2CO_3(aq)}^{2Na^{1+}(aq) + CO_3^{2-}(aq)}$$

$$\longrightarrow Ag_2CO_3(s) + \overbrace{2Na^{1+}(aq) + 2NO_3^{1-}(aq)}^{2NaNO_3(aq)}$$

# Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and $Na_2CO_3$ are mixed.

4. **Net Ionic Equation**: Eliminate the spectator ions from the ionic equation.

$$2Ag^{1+}(aq) + \cancel{2NO_3^{1-}(aq)} + \cancel{2Na^{1+}(aq)} + CO_3^{2-}(aq)$$

$$\longrightarrow Ag_2CO_3(s) + \cancel{2Na^{1+}(aq)} + \cancel{2NO_3^{1-}(aq)}$$

$$\boxed{2Ag^{1+}(aq) + CO_3^{2-}(aq) \longrightarrow Ag_2CO_3(s)}$$

# Acids, Bases, and Neutralization Reactions

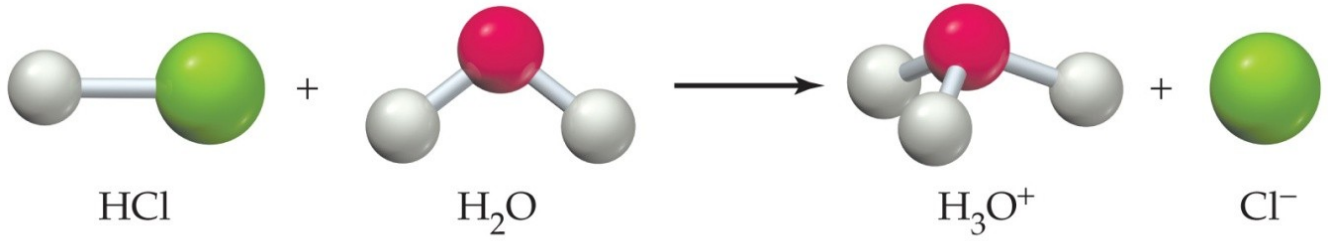
**Arrhenius Acid**: A substance that dissociates in water to produce hydrogen ions, $H^{1+}$:

$$HA(aq) \longrightarrow H^{1+}(aq) + A^{1-}(aq)$$

$$HCl(aq) \longrightarrow H^{1+}(aq) + Cl^{1-}(aq)$$

In water, acids produce hydronium ions, $H_3O^{1+}$:

$$HCl(aq) + H_2O(aq) \longrightarrow H_3O^{1+}(aq) + Cl^{1-}(aq)$$



# Acids, Bases, and Neutralization Reactions

**Arrhenius Base**: A substance that dissociates in water to produce hydroxide ions, $OH^{1-}$:

$$MOH(aq) \longrightarrow M^{1+}(aq) + OH^{1-}(aq)$$

$$NaOH(aq) \longrightarrow Na^{1+}(aq) + OH^{1-}(aq)$$

Ammonia in water, commonly called "ammonium hydroxide", is a base:

$$NH_3(aq) + H_2O(aq) \rightleftharpoons NH_4^{1+}(aq) + OH^{1-}(aq)$$

# Acids, Bases, and Neutralization Reactions

**TABLE 7.2** Some Common Acids and Bases

| | | | | | |
|---|---|---|---|---|---|
| Strong acid | $HClO_4$ | Perchloric acid | $NaOH$ | Sodium hydroxide | Strong base |
| | $H_2SO_4$ | Sulfuric acid | $KOH$ | Potassium hydroxide | |
| | $HBr$ | Hydrobromic acid | $Ba(OH)_2$ | Barium hydroxide | |
| | $HCl$ | Hydrochloric acid | $Ca(OH)_2$ | Calcium hydroxide | |
| | $HNO_3$ | Nitric acid | | | |
| Weak acid | $H_3PO_4$ | Phosphoric acid | $NH_3$ | Ammonia | Weak base |
| | $HF$ | Hydrofluoric acid | | | |
| | $CH_3CO_2H$ | Acetic acid | | | |

**Strong** acids and **strong** bases are **strong** electrolytes.

**Weak** acids and **weak** bases are **weak** electrolytes.

## TABLE 7.3 Common Oxoacids and Their Anions

| Oxoacid | | Oxoanion | |
|---|---|---|---|
| $HNO_2$ | Nitr*ous* acid | $NO_2^-$ | Nitr*ite* ion |
| $HNO_3$ | Nitr*ic* acid | $NO_3^-$ | Nitr*ate* ion |
| $H_3PO_4$ | Phosphor*ic* acid | $PO_4^{3-}$ | Phosph*ate* ion |
| $H_2SO_3$ | Sulfur*ous* acid | $SO_3^{2-}$ | Sulf*ite* ion |
| $H_2SO_4$ | Sulfur*ic* acid | $SO_4^{2-}$ | Sulf*ate* ion |
| HClO | *Hypo*chlor*ous* acid | $ClO^-$ | *Hypo*chlor*ite* ion |
| $HClO_2$ | Chlor*ous* acid | $ClO_2^-$ | Chlor*ite* ion |
| $HClO_3$ | Chlor*ic* acid | $ClO_3^-$ | Chlor*ate* ion |
| $HClO_4$ | *Per*chlor*ic* acid | $ClO_4^-$ | *Per*chlor*ate* ion |

"ous" ⟷ "ite"

"ic" ⟷ "ate"

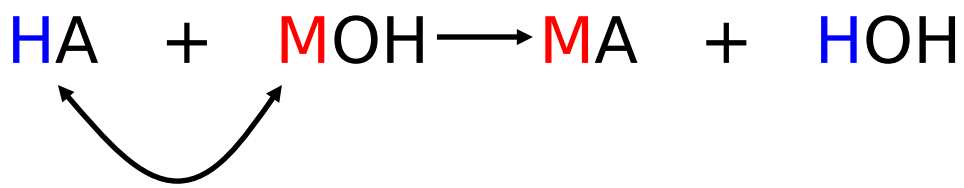
# Acids, Bases, and Neutralization Reactions

## Binary Acids

| | |
|---|---|
| HCl | **hydro**chlor**ic** acid |
| HBr | **hydro**brom**ic** acid |
| HF | **hydro**fluor**ic** acid |

# Acids, Bases, and Neutralization Reactions

These acid-base neutralization reactions are double-replacement reactions just like the precipitation reactions:




or

$$HA + MOH \longrightarrow MA + H_2O$$

Acid Base Salt Water

# Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous $Ba(OH)_2$.

1. Write the chemical formulas of the products (use proper ionic rules for the salt).

$$HBr(aq) + Ba(OH)_2(aq) \longrightarrow H_2O + BaBr_2$$

Acid Base Water Salt

# Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous $Ba(OH)_2$.

2. **Molecular Equation**: Balance the equation and predict the solubility of the salt in the products.

$$2HBr(aq) + Ba(OH)_2(aq) \longrightarrow 2H_2O(l) + \boxed{BaBr_2(aq)}$$

Use the solubility rules.

- **Halides are soluble: $Cl^{1-}$, $Br^{1-}$, $I^{1-}$**
  - ***except* compounds of $Ag^{1+}$, $Hg_2^{2+}$, and $Pb^{2+}$**

# Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous $Ba(OH)_2$.

3. **Ionic Equation**: Dissociate a strong acid and the soluble ionic compounds.

$$\overbrace{2HBr(aq)}^{} \qquad \overbrace{Ba(OH)_2(aq)}^{}$$

$$\overbrace{2H^{1+}(aq) + 2Br^{1-}(aq)}^{} + \overbrace{Ba^{2+}(aq) + 2OH^{1-}(aq)}^{}$$

$$\longrightarrow 2H_2O(l) + \overbrace{Ba^{2+}(aq) + 2Br^{1-}(aq)}^{}$$

$$\overbrace{BaBr_2(aq)}^{}$$

# Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous $Ba(OH)_2$.

4. **Net Ionic Equation**: Eliminate the spectator ions from the ionic equation.

$$2H^{1+}(aq) + \cancel{2Br^{1-}(aq)} + \cancel{Ba^{2+}(aq)} + 2OH^{1-}(aq)$$

$$\longrightarrow 2H_2O(l) + \cancel{Ba^{2+}(aq)} + \cancel{2Br^{1-}(aq)}$$

$$2H^{1+}(aq) + 2OH^{1-}(aq) \longrightarrow 2H_2O(l)$$

or

$$\boxed{H^{1+}(aq) + OH^{1-}(aq) \longrightarrow H_2O(l)}$$

# Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

1. Write the chemical formulas of the products (use proper ionic rules for the salt).

$$HF(aq) + NaOH(aq) \longrightarrow H_2O + NaF$$

Acid Base Water Salt

# Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

2. **Molecular Equation**: Balance the equation and predict the solubility of the salt in the products.

$$HF(aq) + NaOH(aq) \longrightarrow H_2O(l) + \boxed{NaF(aq)}$$

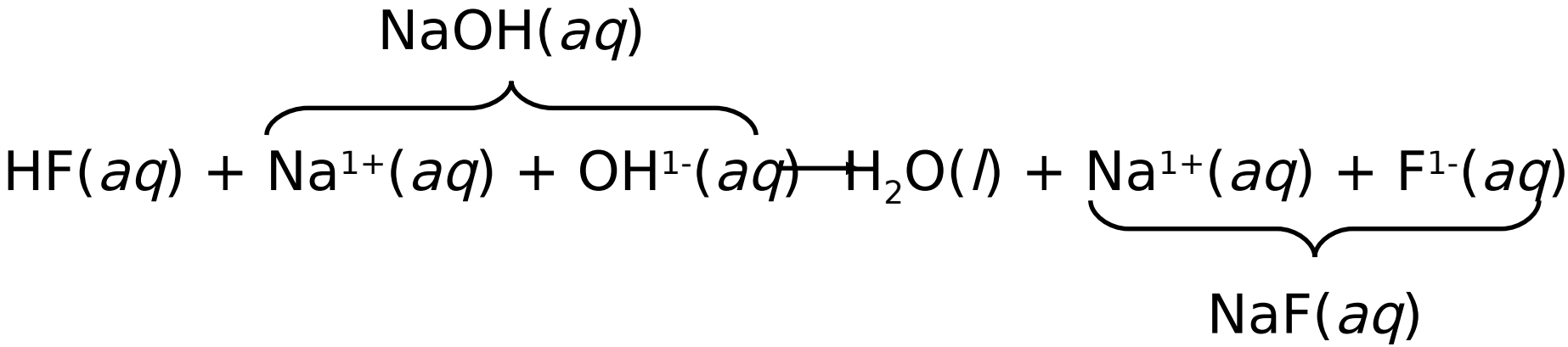
Use the solubility rules.

- **Compounds based on Group 1A cations: $Li^{1+}$, $Na^{1+}$, $K^{1+}$, $Cs^{1+}$ are soluble.**

# Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

3. **Ionic Equation**: Dissociate a *strong* acid and the soluble ionic compounds.



# Acids, Bases, and Neutralization Reactions

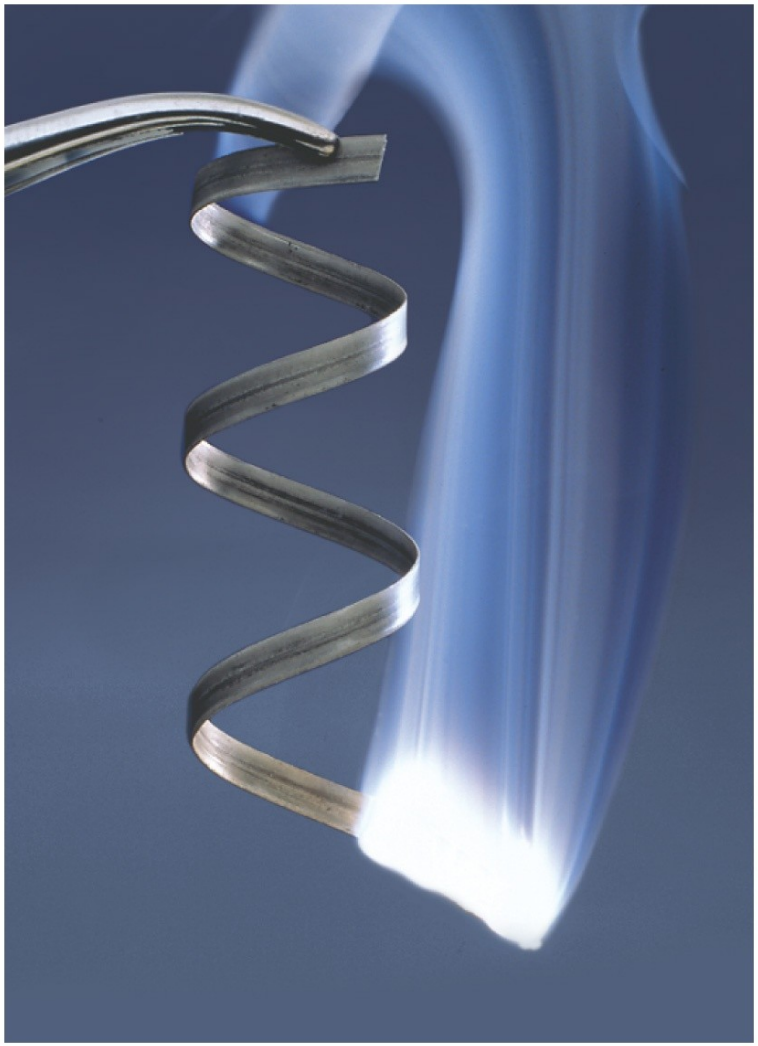
Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

4. **Net Ionic Equation**: Eliminate the spectator ions from the ionic equation.

$$HF(aq) + \cancel{Na^{1+}(aq)} + OH^{1-}(aq) \longrightarrow H_2O(l) + \cancel{Na^{1+}(aq)} + F^{1-}(aq)$$

$$\boxed{HF(aq) + OH^{1-}(aq) \longrightarrow H_2O(l) + F^{1-}(aq)}$$

# Oxidation-Reduction Reactions (Redox)

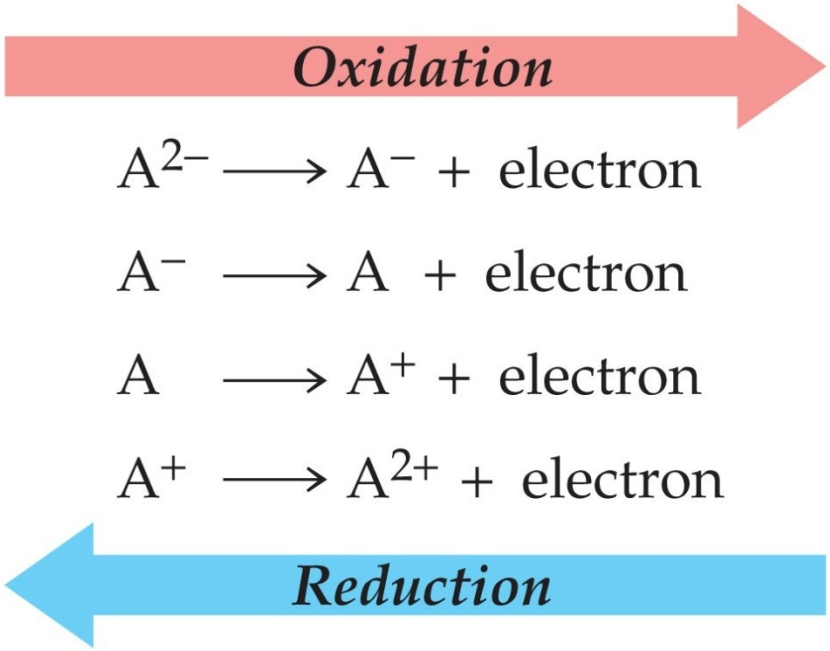


# Redox Reactions - OIL RIG

## (oxidation is loss / reduction is gain)

$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$

Rusting of iron: an **oxidation rxn** of Fe (Fe loses $e^-$)




Reactant A might be:
a neutral atom,
a monatomic ion,
a polyatomic ion,
or a molecule.

$2Fe_2O_3(s) + 3C(s) \longrightarrow 4Fe(s) + 3CO_2(g)$

Manufacture of iron: a **reduction rxn** of Fe (Fe gains $e^-$)

# Oxidation-Reduction (Redox) Reactions

## OIL RIG

**Oxidation**: The **loss** of one or more electrons by a substance, whether element, compound, or ion.

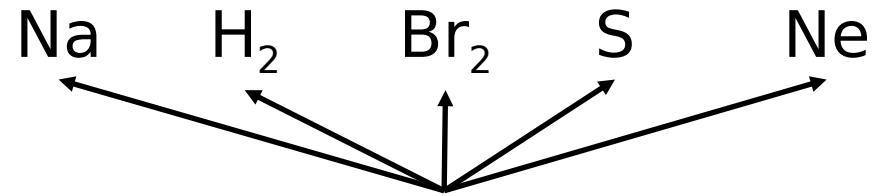
**Reduction**: The **gain** of one or more electrons by a substance, whether element, compound, or ion.

# Oxidation-Reduction (Redox) Reactions

**Oxidation Number (State)**: A value which indicates whether an atom is neutral, electron-rich, or electron-poor.

## Rules for Assigning Oxidation Numbers

1. An atom in its elemental state has an oxidation number of 0.



# Oxidation-Reduction (Redox) Reactions

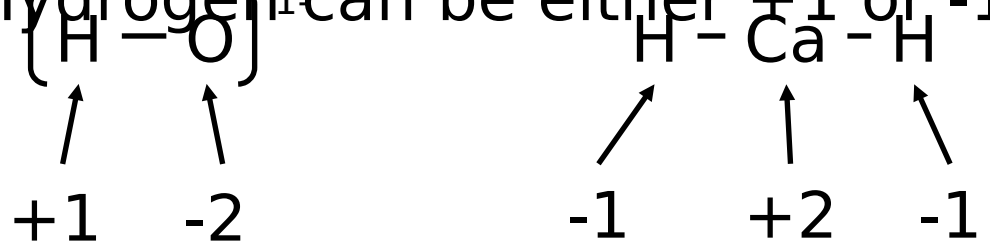
2. An atom in a monatomic ion has an oxidation number identical to its charge.

$Na^{1+}$ $Ca^{2+}$ $Al^{3+}$ $Cl^{1-}$ $O^{2-}$

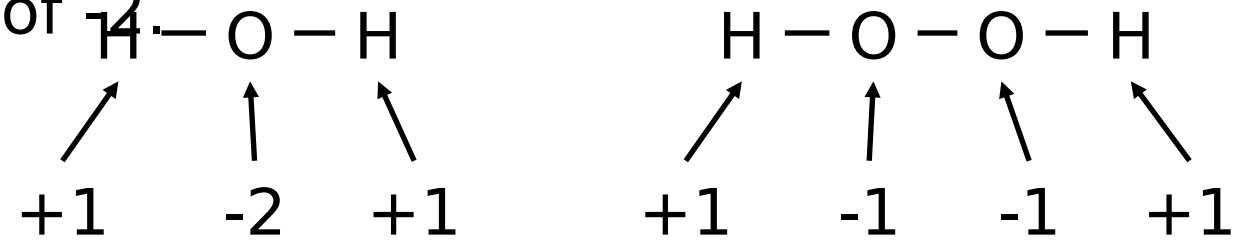
+1 +2 +3 -1 -2

# Oxidation-Reduction (Redox) Reactions

3. An atom in a polyatomic ion or in a molecular compound usually has the same oxidation number it would have if it were a monatomic ion.
   a) Hydrogen can be either +1 or -1.

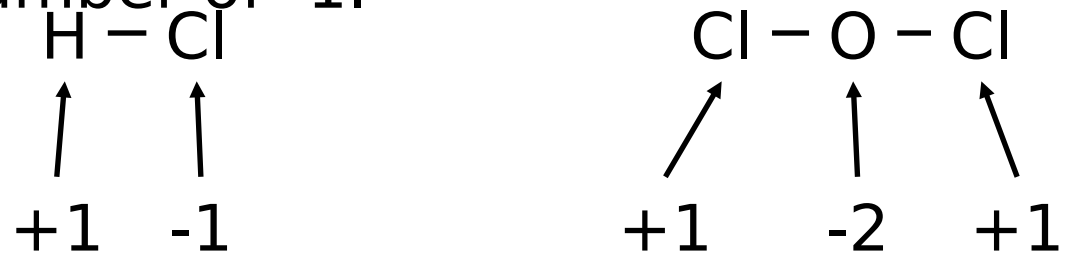



   b) Oxygen *usually* has an oxidation number of -2.



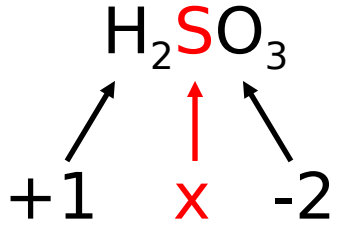
# Oxidation-Reduction (Redox) Reactions

3. c) Halogens *usually* have an oxidation number of -1.



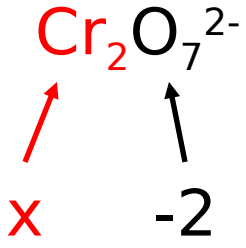
# Oxidation-Reduction (Redox) Reactions

4. The sum of the oxidation numbers is 0 for a neutral compound and is equal to the net charge for a polyatomic ion.




$2(+1) + x + 3(-2) = 0$ (net charge)

$x = +4$




$2(x) + 7(-2) = -2$ (net charge)

$x = +6$

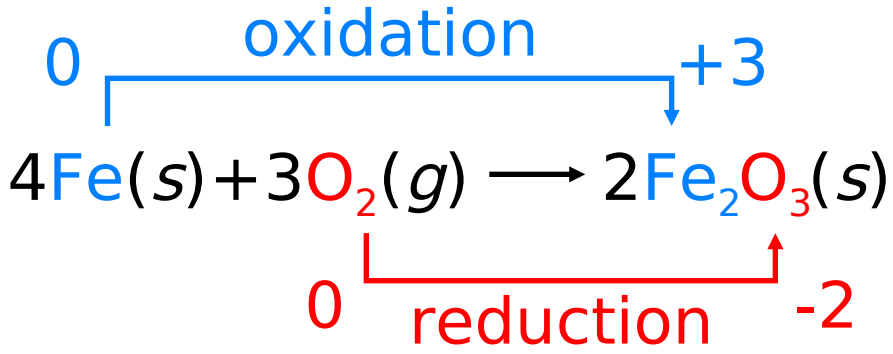
# Identifying Redox Reactions

## Reducing Agent (Fe)

- Causes reduction
- Loses one or more $e^-$
- Undergoes oxidation
- Atom's oxidation # increases

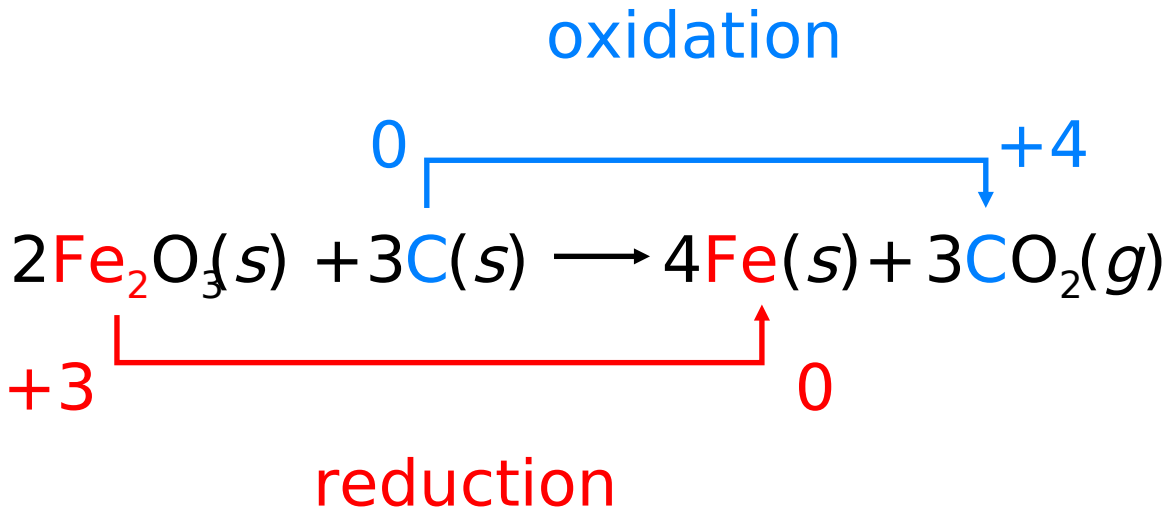
## Oxidizing Agent ($O_2$)

- Causes oxidation
- Gains one or more $e^-$
- Undergoes reduction
- Atom's oxidation # decreases



# Identifying Redox Reactions

**Reducing Agent (C)**




**Oxidizing Agent (Fe)**

# Identify the Oxidizing and Reducing Agents in Each of the Following

$3\ H_2S + 2\ NO_3^- + 2\ H^+ \rightarrow 3\ S + 2\ NO + 4\ H_2O$

$MnO_2 + 4\ HBr \rightarrow MnBr_2 + Br_2 + 2\ H_2O$

# Identify the Oxidizing and Reducing Agents in Each of the Following

red agent ox agent

$$3\ H_2S + 2\ NO_3^- + 2\ H^+ \rightarrow 3\ S + 2\ NO + 4\ H_2O$$

-2 +5 -2 +1 0 +2 -2 +1

oxidation

reduction

ox ag red ag

$$MnO_2 + 4\ HBr \rightarrow MnBr_2 + Br_2 + 2\ H_2O$$

+4 -2 +1 -1 +2 -1 0 +1 -2

oxidation

reduction

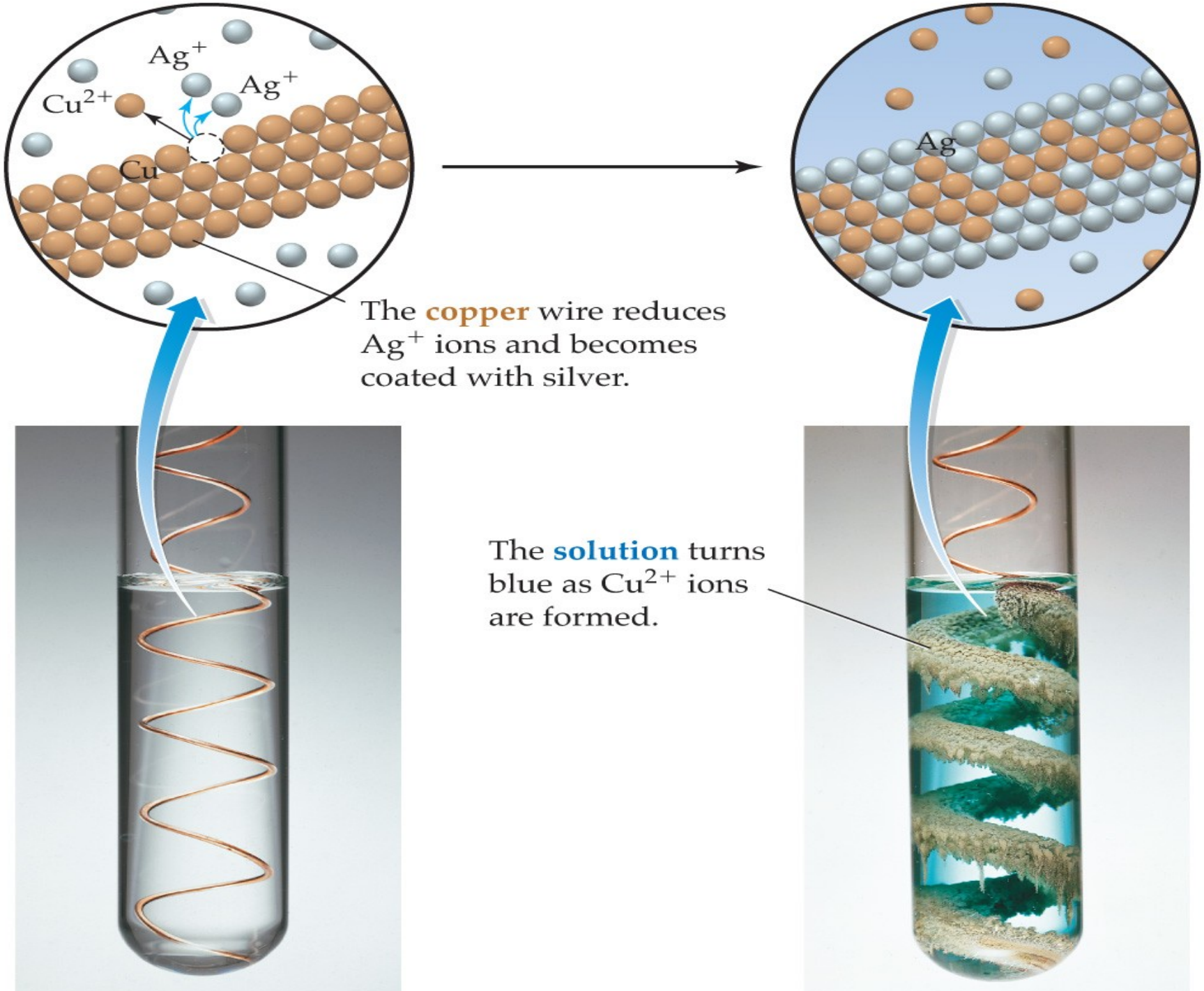
# The redox Activity Series of the Elements

## Which one of these reactions will occur?

**The forward rxn or the reverse rxn?**

$$Cu(s) + 2Ag^{1+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$$

$$2Ag(s) + Cu^{2+}(aq) \longrightarrow 2Ag^{1+}(aq) + Cu(s)$$


Ag+
Ag+
Cu2+
Cu
Ag
The copper wire reduces Ag+ ions and becomes coated with silver.
The solution turns blue as Cu2+ ions are formed.

# The Activity Series of the Elements

**TABLE 7.4** A Partial Activity Series of the Elements

| Oxidation Reaction | | |
|---|---|---|
| Strongly reducing | $Li \rightarrow Li^+ + e^-$ | These elements react rapidly with aqueous $H^+$ ions (acid) or with liquid $H_2O$ to release $H_2$ gas. |
| | $K \rightarrow K^+ + e^-$ | |
| | $Ba \rightarrow Ba^{2+} + 2\,e^-$ | |
| | $Ca \rightarrow Ca^{2+} + 2\,e^-$ | |
| | $Na \rightarrow Na^+ + e^-$ | |
| | $Mg \rightarrow Mg^{2+} + 2\,e^-$ | These elements react with aqueous $H^+$ ions or with steam to release $H_2$ gas. |
| | $Al \rightarrow Al^{3+} + 3\,e^-$ | |
| | $Mn \rightarrow Mn^{2+} + 2\,e^-$ | |
| | $Zn \rightarrow Zn^{2+} + 2\,e^-$ | |
| | $Cr \rightarrow Cr^{3+} + 3\,e^-$ | |
| | $Fe \rightarrow Fe^{2+} + 2\,e^-$ | |
| | $Co \rightarrow Co^{2+} + 2\,e^-$ | These elements react with aqueous $H^+$ ions to release $H_2$ gas. |
| | $Ni \rightarrow Ni^{2+} + 2\,e^-$ | |
| | $Sn \rightarrow Sn^{2+} + 2\,e^-$ | |
| | $H_2 \rightarrow 2\,H^+ + 2\,e^-$ | |
| | $Cu \rightarrow Cu^{2+} + 2\,e^-$ | These elements do not react with aqueous $H^+$ ions to release $H_2$. |
| | $Ag \rightarrow Ag^+ + e^-$ | |
| | $Hg \rightarrow Hg^{2+} + 2\,e^-$ | |
| | $Pt \rightarrow Pt^{2+} + 2\,e^-$ | |
| Weakly reducing | $Au \rightarrow Au^{3+} + 3\,e^-$ | |

The elements that are higher up in the table are more likely to be oxidized.

Thus, any element higher in the activity series will reduce the ion of any element lower in the activity series.

## TABLE 18.1 Standard Reduction Potentials at 25 °C

Stronger oxidizing agent ↑

| Reduction Half-Reaction | $E°$ (V) |
|---|---|
| $F_2(g) + 2\,e^- \longrightarrow 2\,F^-(aq)$ | 2.87 |
| $H_2O_2(aq) + 2\,H^+(aq) + 2\,e^- \longrightarrow 2\,H_2O(l)$ | 1.78 |
| $PbO_2(s) + 4\,H^+(aq) + SO_4^{2-}(aq) + 2\,e^- \longrightarrow PbSO_4(s) + 2\,H_2O(l)$ | 1.69 |
| $MnO_4^-(aq) + 4\,H^+(aq) + 3\,e^- \longrightarrow MnO_2(s) + 2\,H_2O(l)$ | 1.68 |
| $MnO_4^-(aq) + 8\,H^+(aq) + 5\,e^- \longrightarrow Mn^{2+}(aq) + 4\,H_2O(l)$ | 1.51 |
| $Au^{3+}(aq) + 3\,e^- \longrightarrow Au(s)$ | 1.50 |
| $PbO_2(s) + 4\,H^+(aq) + 2\,e^- \longrightarrow Pb^{2+}(aq) + 2\,H_2O(l)$ | 1.46 |
| $Cl_2(g) + 2\,e^- \longrightarrow 2\,Cl^-(aq)$ | 1.36 |
| $Cr_2O_7^{2-}(aq) + 14\,H^+(aq) + 6\,e^- \longrightarrow 2\,Cr^{3+}(aq) + 7\,H_2O(l)$ | 1.33 |
| $O_2(g) + 4\,H^+(aq) + 4\,e^- \longrightarrow 2\,H_2O(l)$ | 1.23 |
| $MnO_2(s) + 4\,H^+(aq) + 2\,e^- \longrightarrow Mn^{2+}(aq) + 2\,H_2O(l)$ | 1.21 |
| $IO_3^-(aq) + 6\,H^+(aq) + 5\,e^- \longrightarrow \frac{1}{2}\,I_2(aq) + 3\,H_2O(l)$ | 1.20 |
| $VO_2^+(aq) + 2\,H^+(aq) + e^- \longrightarrow VO^{2+}(aq) + H_2O(l)$ | 1.00 |
| $Br_2(l) + 2\,e^- \longrightarrow 2\,Br^-(aq)$ | 1.09 |
| $NO_3^-(aq) + 4\,H^+(aq) + 3\,e^- \longrightarrow NO(s) + 2\,H_2O(l)$ | 0.96 |
| $ClO_2(g) + e^- \longrightarrow ClO_2^-(aq)$ | 0.95 |
| $Ag^+(aq) + e^- \longrightarrow Ag(s)$ | 0.80 |
| $Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$ | 0.77 |
| $O_2(g) + 2\,H^+(aq) + 2\,e^- \longrightarrow H_2O_2(aq)$ | 0.70 |
| $MnO_4^-(aq) + e^- \longrightarrow MnO_4^{2-}(aq)$ | 0.56 |
| $I_2(s) + 2\,e^- \longrightarrow 2\,I^-(aq)$ | 0.54 |
| $Cu^+(aq) + e^- \longrightarrow Cu(s)$ | 0.52 |
| $O_2(g) + 2\,H_2O(l) + 4\,e^- \longrightarrow 4\,OH^-(aq)$ | 0.40 |

Weaker reducing agent ↓

| | | |
|---|---|---|
| $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$ | | 0.34 |
| $SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \longrightarrow H_2SO_3(aq) + H_2O(l)$ | | 0.20 |
| $Cu^{2+}(aq) + e^- \longrightarrow Cu^+(aq)$ | | 0.16 |
| $Sn^{4+}(aq) + 2e^- \longrightarrow Sn^{2+}(aq)$ | | 0.15 |
| $2H^+(aq) + 2e^- \longrightarrow H_2(s)$ | | 0 |
| $Fe^{3+}(aq) + 3e^- \longrightarrow Fe(s)$ | | −0.036 |
| $Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$ | | −0.13 |
| $Sn^{2+}(aq) + 2e^- \longrightarrow Sn(s)$ | | −0.14 |
| $Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$ | | −0.23 |
| $Cd^{2+}(aq) + 2e^- \longrightarrow Cd(s)$ | | −0.40 |
| $Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$ | | −0.45 |
| $Cr^{3+}(aq) + e^- \longrightarrow Cr^{2+}(aq)$ | | −0.50 |
| $Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$ | | −0.76 |
| $Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$ | | −0.73 |
| $2H_2O(l) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)$ | | −0.83 |
| $Mn^{2+}(aq) + 2e^- \longrightarrow Mn(s)$ | | −1.18 |
| $Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$ | | −1.66 |
| $Mg^{2+}(aq) + 2e^- \longrightarrow Mg(s)$ | | −2.37 |
| $Na^+(aq) + e^- \longrightarrow Na(s)$ | | −2.71 |
| $Ca^{2+}(aq) + 2e^- \longrightarrow Ca(s)$ | | −2.76 |
| $Ba^{2+}(aq) + 2e^- \longrightarrow Ba(s)$ | | −2.90 |
| $K^+(aq) + e^- \longrightarrow K(s)$ | | −2.92 |
| $Li^+(aq) + e^- \longrightarrow Li(s)$ | | −3.04 |

Weaker oxidizing agent

Stronger reducing agent

# Half Reactions

**Given a total reaction**

$$Cu(s) + Ag^{1+}(aq) \longrightarrow Cu^{2+}(aq) + Ag(s)$$

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Isolate the reaction of Cu in one equation

$$Cu(s) \longrightarrow Cu^{2+}(aq)$$

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Isolate the reaction of Ag in one equation

$$Ag^{1+}(g) \longrightarrow Ag(s)$$

Find the correct half reaction on the following lists

# Balancing Redox Reactions:

## Basic Solution

## Acidic Solution

# The Half-Reaction Method

# Balancing Redox Rxns (Basic Solution)

1) Write the unbalanced equation
2) Assign oxidation numbers to the atoms
   a) determine which element is oxidized and which is reduced
3) Write redox half-reactions, including electrons
   1) oxidation :electrons on the right of arrow
   2) reduction: electrons on left of arrow
4) Balance the half-reactions by mass
   **1) first balance elements other than H and O**
   **2) next, balance O by adding $H_2O$ wherever an O is needed**
   **3) then add $H^{+1}$ where a H is needed**
   **4) neutralize all $H^+$ (acid) with $OH^-$ (base)**

a) Balance half-reactions by charge
   a) balance charge by adjusting electrons

1) Balance electrons between half-reactions
2) Add half-reactions together
3) Check the answer

# Balancing Redox Rxns (Acid Solution)

1) Write the unbalanced equation
2) Assign oxidation numbers to the atoms
   a) determine which element is oxidized and which is reduced
3) Write redox half-reactions, including electrons
   1) oxidation :electrons on the right of arrow
   2) reduction: electrons on left of arrow
4) Balance the half-reactions by mass
   **a) first balance elements other than H and O**
   **b) next, balance O by adding $H_2O$ wherever an O is needed**
   **c) then add $H^{+1}$ where a H is needed**
5) Balance half-reactions by charge
   1) balance charge by adjusting electrons

a) Balance electrons between half-reactions

b) Add half-reactions together

c) Check the answer

# Balancing Redox Reactions: The Half-Reaction Method

Balance the following net ionic equation in acidic solution:

STEP 1: WRITE THE UNBALANCED EQUATION

$$I^{1-}(aq) + Cr_2O_7^{2-}(aq) \longrightarrow Cr^{3+}(aq) + IO_3^{1-}(aq)$$

STEP 2: ASSIGN OXIDATION STATES

$$I^{1-}(aq) + Cr_2O_7^{2-}(aq) \longrightarrow Cr^{3+}(aq) + IO_3^{1-}(aq)$$

1- (6+)(2-) 3+ (5+) (2-)

I goes from 1- to 5+ (LOSES e-) = OXIDIZED

Cr goes from 6+ to 3+ (GAINS e-) = REDUCED

# Balancing Redox Reactions: The Half-Reaction Method

STEP 3: Write the two unbalanced half-reactions

OX: $I^{1-}(aq) \longrightarrow IO_3^{1-}(aq) +$

RED: $e^- + Cr_2O_7^{2-}(aq) \longrightarrow Cr^{3+}(aq)$

EP 4: Balance the half-reactions by mass (excluding O & H)

OX: $I^{1-}(aq) \longrightarrow IO_3^{1-}(aq)$

one Iodine ⟶ one Iodine

RED: $Cr_2O_7^{2-}(aq) \longrightarrow 2Cr^{3+}(aq)$

two Chromium ⟶ two Chromium

# Balancing Redox Reactions: The Half-Reaction Method

STEP 4: Balance the half-reactions by mass – add water wherever O is needed.

OX: $I^{1-}(aq) + 3H_2O \longrightarrow IO_3^{1-}(aq)$

RED: $Cr_2O_7^{2-}(aq) \longrightarrow 2Cr^{3+}(aq) + 7H_2O$

STEP 4: Balance the half-reactions by mass – add $H^+$ wherever H is needed.

OX: $I^{1-}(aq) + 3H_2O \longrightarrow IO_3^{1-}(aq) + 6H^+$

RED: $Cr_2O_7^{2-}(aq) + 14H^+ \longrightarrow 2Cr^{3+}(aq) + 7H_2O$

# Balancing Redox Reactions: The Half-Reaction Method

STEP 5: Balance the half-reactions by charge

$I^{1-}(aq) + 3H_2O \longrightarrow IO_3^{1-}(aq) + 6H^+ + x\,e^-$

$(1-)=1-$ $\quad (1-)+(6+)+(x)=1-$ $\quad X=6-$

$Cr_2O_7^{2-}(aq) + 14H^+ + x\,e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O$

$(2-)+(14+)+(x)=6+$ $\quad X=6-$ $\quad (6+)=6+$

# Balancing Redox Reactions: The Half-Reaction Method

STEP 6 : Balance electrons between half-reactions

$$I^{1-}(aq) + 3H_2O \longrightarrow IO_3^{1-}(aq) + 6H^+ + 6e^-$$

$$Cr_2O_7^{2-}(aq) + 14H^+ + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O$$

Electrons are balanced as is

STEP 7 : Add the half-reactions

$$I^{1-}(aq) + 3H_2O \longrightarrow IO_3^{1-}(aq) + 6H^+ + 6e^-$$

$$Cr_2O_7^{2-}(aq) + 14H^+ + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O$$

$\times 4$ and $\times 8$

$$I^{1-}(aq) + Cr_2O_7^{2-}(aq) + 8H^+ \longrightarrow IO_3^{1-}(aq) + 2Cr^{3+}(aq) + 4H_2O$$

# Balancing Redox Reactions: The Half-Reaction Method

Balance the following net ionic equation in basic solution:

$$MnO_4^{1-}(aq) + Br^{1-}(aq) \longrightarrow MnO_2(s) + BrO_3^{1-}(aq)$$

# Balancing Redox Reactions: The Half-Reaction Method

- Write the two unbalanced half-reactions.

Ox: $Br^{1-}(aq) \longrightarrow BrO_3^{1-}(aq)$

Red: $MnO_4^{1-}(aq) \longrightarrow MnO_2(s)$

- Write oxidation states

Br- = -1 $\xrightarrow[-6\ e-]{}$ BrO3- X+ 3(-2)=1-
X=+5

$MnO_4^{1-}$ X+ 4(-2)=1- $\xrightarrow[+3\ e-]{}$ $MnO_2$ X+ 2(-2)=0
X=+7 X=+4

# Balancing Redox Reactions: The Half-Reaction Method

- Balance both half-reactions for all atoms except O and H.

$$Br^{1-}(aq) \longrightarrow BrO_3^{1-}(aq) + e^-$$

$$e^- + MnO_4^{1-}(aq) \longrightarrow MnO_2(s)$$

  - ❑ Both Br and Mn are balanced

- Balance O by adding H2O

$$3H_2O + Br^{1-}(aq) \longrightarrow BrO_3^{1-}(aq)$$

$$MnO_4^{1-}(aq) \longrightarrow MnO_2(s) + 2H_2O$$

# Balancing Redox Reactions: The Half-Reaction Method

- Balance each half-reaction for H by adding $H^{1+}$.

$$3H_2O(l) + Br^{1-}(aq) \longrightarrow BrO_3^{1-}(aq) + 6H^{1+}(aq)$$

$$+4H^{1+}(aq) + MnO_4^{1-}(aq) \longrightarrow MnO_2(s) + 2H_2O(l)$$

- Since in basic solution, add $OH^-$ for each $H^{1+}$.

$$3H_2O(l) + Br^{1-}(aq) + 6OH^- \longrightarrow BrO_3^{1-}(aq) + 6H^{1+} + 6OH^-$$

$$3H_2O(l) + Br^{1-}(aq) + \longrightarrow BrO_3^{1-}(aq) + 6H_2O \quad 3$$

$$6OH^- \quad 4OH^- + 4H^{1+} + MnO_4^{1-}(aq) \longrightarrow MnO_2(s) + 2H_2O(l) + 4OH^-$$

$$4H_2O(l) + MnO_4^{1-}(aq) \longrightarrow MnO_2(s) + 2H_2O(l) + 4OH^-$$

$$2$$

# Balancing Redox Reactions: The Half-Reaction Method

- Balance each half-reaction for charge by adding electrons to the side with greater positive charge.

$Br^{1-}(aq) + 6OH^- \longrightarrow BrO_3^{1-}(aq) + 3H_2O(l) + 6\ e^-$

(1-)+(6-)=7- (1-)+(x)=7- x=6-

$2e^- + 2H_2O(l) + MnO_4^{1-}(aq) \longrightarrow MnO_2(s) + 4OH^-$

x+(1-)=4- X=3- 4-

# Balancing Redox Reactions: The Half-Reaction Method

- Multiply each half-reaction by a factor to make the electron count the same in both half-reactions.

$$Br^{1-}(aq) + 6OH^- \longrightarrow BrO_3^{1-}(aq) + 3H_2O(l) + 6e^-$$

$$2\left[3e^- + 2H_2O(l) + MnO_4^{1-}(aq) \longrightarrow MnO_2(s) + 4OH^-\right]$$

# Balancing Redox Reactions: The Half-Reaction Method

- Add the two balanced half-reactions together and cancel species that appear on both sides of the equation.

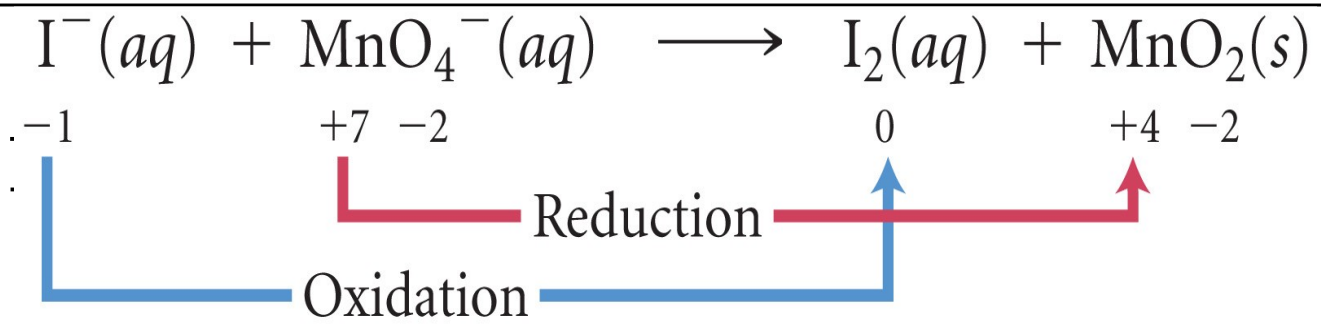
$$6OH^-(aq) + Br^{1-}(aq) \longrightarrow BrO_3^{1-}(aq) + 3H_2O(l) + 6e^-$$

$$6e^- + 4H_2O(l) + 2MnO_4^{1-}(aq) \longrightarrow 2MnO_2(s) + 8OH^-(aq)$$

$$\times 2$$

$$H_2O(l) + 2MnO_4^{1-}(aq) + Br^{1-}(aq) \longrightarrow 2MnO_2(s) + 2OH^-(aq) + BrO_3^{1-}(aq)$$

# Balance the equation:
# $I^-_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow I_{2(aq)} + MnO_{2(s)}$ in basic solution

| | |
|---|---|
| Assign Oxidation States |   |
| Separate into half-reactions | ox: $I^-_{(aq)} \rightarrow I_{2(aq)}$ red: $MnO_4^-{}_{(aq)} \rightarrow MnO_{2(s)}$ |

# Balance the equation: $I^-_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow I_{2(aq)} + MnO_{2(s)}$ in basic solution

| | |
|---|---|
| Balance half-reactions by mass | ox: $I^-_{(aq)} \rightarrow I_{2(aq)}$<br>red: $MnO_4^-{}_{(aq)} \rightarrow MnO_{2(s)}$ |
| Balance half-reactions by mass | ox: $2\ I^-_{(aq)} \rightarrow I_{2(aq)}$<br>red: $MnO_4^-{}_{(aq)} \rightarrow MnO_{2(s)}$ |
| Balance half-reactions by mass, then O by adding | ox: $2\ I^-_{(aq)} \rightarrow I_{2(aq)}$<br>red: $MnO_4^-{}_{(aq)} \rightarrow MnO_{2(s)} + 2\ H_2O_{(l)}$ |

# Balance the equation: $I^-_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow I_{2(aq)} + MnO_{2(s)}$ in basic solution

| | |
|---|---|
| Balance half-rxns by mass then H by | ox: $2\ I^-_{(aq)} \rightarrow I_{2(aq)}$<br>red: $4\ H^+_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow MnO_{2(s)} + 2\ H_2O_{(l)}$ |
| Balance by adding $H^+$ | ox: $2\ I^-_{(aq)} \rightarrow I_{2(aq)}$<br>red: $4\ H^+_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow MnO_{2(s)} + 2\ H_2O_{(l)}$ |
| half-rxns by mass in base, neutralize the | $4\ H^+_{(aq)} + \mathbf{4\ OH^-_{(aq)}} + MnO_4^-{}_{(aq)} \rightarrow MnO_{2(s)} + 2\ H_2O_{(l)} + \mathbf{4\ OH^-_{(aq)}}$<br>$2\ H_2O$ |

# Balance the equation: $I^-_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow I_{2(aq)} + MnO_{2(s)}$ in basic solution

| | |
|---|---|
| Balance Half-rxns by charge | ox: $2\ I^-_{(aq)} \rightarrow I_{2(aq)} + 2\ e^-$<br>red: $MnO_4^-{}_{(aq)} + 2\ H_2O_{(l)} + 3\ e^- \rightarrow MnO_{2(s)} + 4\ OH^-_{(aq)}$ |
| Balance electrons between half-rxns | ox: $2\ I^-_{(aq)} \rightarrow I_{2(aq)} + 2\ e^-$ } x3<br>red: $MnO_4^-{}_{(aq)} + 2\ H_2O_{(l)} + 3\ e^- \rightarrow MnO_{2(s)} + 4\ OH^-_{(aq)}$ }x2<br>ox: $6\ I^-_{(aq)} \rightarrow 3\ I_{2(aq)} + 6\ e^-$<br>red: $2\ MnO_4^-{}_{(aq)} + 4\ H_2O_{(l)} + 6\ e^- \rightarrow 2\ MnO_{2(s)} + 8\ OH^-_{(aq)}$ |

# Balance the equation: $I^-_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow I_{2(aq)} + MnO_{2(s)}$ in basic solution

**Add the Half-reactions**

ox: $6\ I^-_{(aq)} \rightarrow 3\ I_{2(aq)} + 6\ e^-$

red: $2\ MnO_4^-{}_{(aq)} + 4\ H_2O_{(l)} + 6\ e^- \rightarrow 2\ MnO_{2(s)} + 8\ OH^-_{(aq)}$

$6\ I^-_{(aq)} + 2\ MnO_4^-{}_{(aq)} + 4\ H_2O_{(l)} \rightarrow 3\ I_{2(aq)} + 2\ MnO_{2(s)} + 8\ OH^-_{(aq)}$

**Check**

| Reactant Count | Element | Product Count |
|---|---|---|
| 6 | I | 6 |
| 2 | Mn | 2 |
| 12 | O | 12 |
| 8 | H | 8 |
| 8- | charge | 8- |

# Practice Balancing Equations

**$H_2O_2 + KI + H_2SO_4 \rightarrow K_2SO_4 + I_2 + H_2O$**

+1 -1 +1 -1 +1 +6 -2 +1 +6 -2 0 +1 -2

oxidation

reduction

ox: $2\ I^{-1} \rightarrow I_2 + 2e^{-1}$

red: $H_2O_2 + 2e^{-1} + 2\ H^+ \rightarrow 2\ H_2O$

tot $2\ I^{-1} + H_2O_2 + 2\ H^+ \rightarrow I_2 + 2\ H_2O$

**$H_2O_2 + 2\ KI + H_2SO_4 \rightarrow K_2SO_4 + 1\ I_2 + 2\ H_2$**

# Practice - Balance the Equation

**$ClO_3^{-1} + Cl^{-1} \rightarrow Cl_2$ (in acid)**

ox: $2\ Cl^{-1} \rightarrow Cl_2 + 2\ e^{-1}$ } x5

red: $2\ ClO_3^{-1} + 10\ e^{-1} + 12\ H^{+} \rightarrow Cl_2 + 6\ H_2O$} x1

tot $10\ Cl^{-1} + 2\ ClO_3^{-1} + 12\ H^{+} \rightarrow 6\ Cl_2 + 6\ H_2O$

**$1\ ClO_3^{-1} + 5\ Cl^{-1} + 6\ H^{+1} \rightarrow 3\ Cl_2 + 3\ H_2O$**

# Redox Stoichiometry

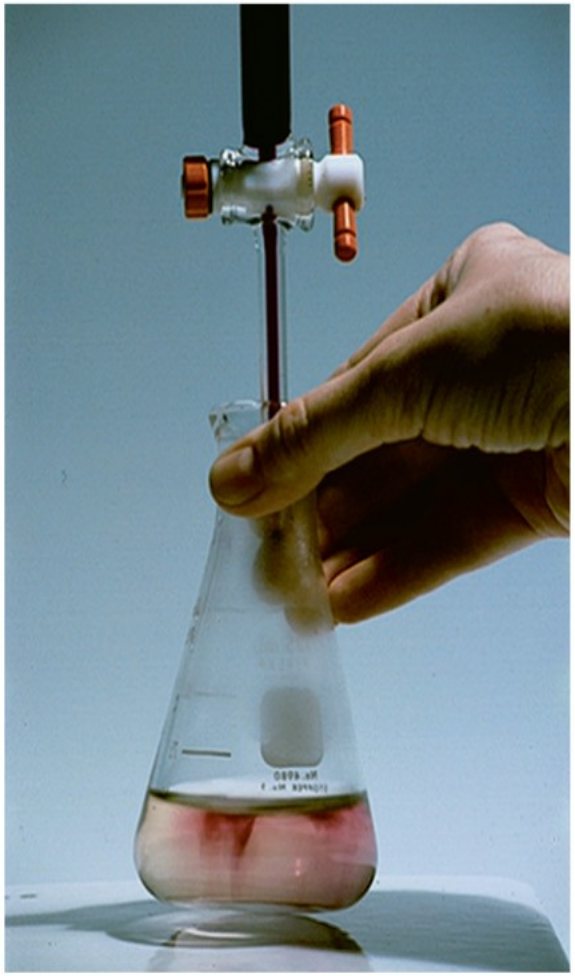
**Titration**: A procedure for determining the concentration of a solution by allowing a carefully measured volume to react with a solution of another substance (the standard solution) whose concentration is known.

(Oxalic acid) + Potassium permanganate

$$5H_2C_2O_4(aq) + 2MnO_4^{1-}(aq) + 6H^{1+}(aq) \longrightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$

$K^+$ is a spectator ion and it is not shown in the net ionic equation

A precise amount of oxalic acid is weighed and dissolved in aqueous $H_2SO_4$.



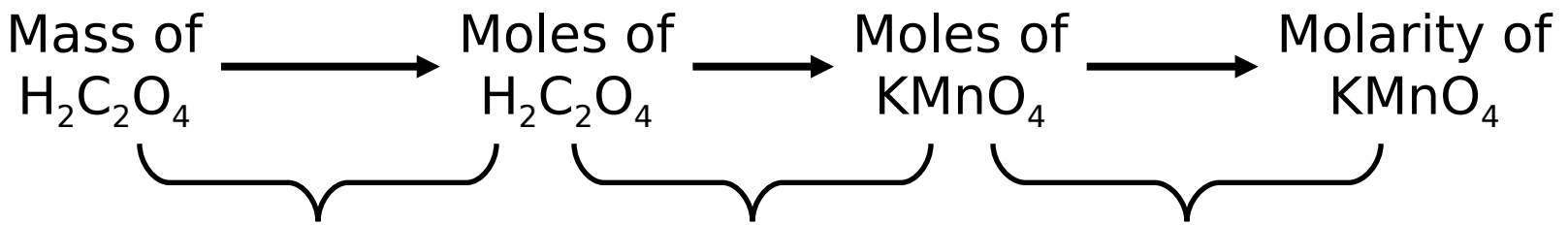
Aqueous $KMnO_4$ of unknown concentration is added from a buret until …

…the purple color persists, indicating that all of the oxalic acid has reacted.

# Redox Stoichiometry

A solution is prepared with 0.2585 g of oxalic acid ($H_2C_2O_4$). 22.35 mL of an unknown potassium permanganate solution are needed to titrate the oxalic acid solution. What is the concentration of the potassium permanganate solution?

$$5H_2C_2O_4(aq) + 2MnO_4^{1-}(aq) + 6H^{1+}(aq) \longrightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$




Molar Mass of $H_2C_2O_4$

$$\frac{1\ \text{mol}}{90.04\ \text{g}}$$

oxalic acid

Mole Ratio

$$\frac{2\ \text{mol}\ KMnO_4}{5\ \text{mol}\ H_2C_2O_4}$$

Molarity of $KMnO_4$

$$\frac{158\ \text{g}\ KMnO_4}{1\ \text{mole}\ KMnO_4}$$

# Redox Stoichiometry

$$5H_2C_2O_4(aq) + 2MnO_4^{1-}(aq) + 6H^{1+}(aq) \longrightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$

**Moles of $H_2C_2O_4$ available**:

$$0.2585 \text{ g } H_2C_2O_4 \times \frac{1 \text{ mol}}{90.04 \text{ g}} = 0.002871 \text{ mol } H_2C_2O_4$$

**Moles of $KMnO_4$ reacted**:

$$0.002871 \text{ mol } H_2C_2O_4 \times \frac{2 \text{ mol } KMnO_4}{5 \text{ mol } H_2C_2O_4} = 0.001148 \text{ mol } KMnO_4$$

# Redox Stoichiometry

$5H_2C_2O_4(aq) + 2MnO_4^{1-}(aq) + 6H^{1+}(aq)$

$\longrightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$

**Concentration of $KMnO_4$ solution**:

$$\frac{0.001148 \text{ mol } KMnO_4}{22.35 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.05136 \text{ M } KMnO_4$$